DOCUMENT RESUME

ED 093 655 SE 017 612

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TITLE	Dynamic Nature of Atoms and Molecules, Science (Experimental): 5316.06.
INSTITUTION	Dade County Public Schools, Miami, Fla.
PUB DATE	72
NOTE	26p.; An Authorized Course of Instruction for the Quinmester Program
EDRS PRICE	MF-$0.75 HC-$1.85 PLUS POSTAGE
DESCRIPTORS	Atomic Theory; *Chemical Equilibrium; *Chemistry; Instruction; *Instructional Materials; Kinetic Molecular Theory; Science Education; *Secondary School Science; *Teaching Guides; Units of Study (Subject Fields)
IDENTIFIERS	*Quinmester Program

ABSTRACT

This unit of instruction deals with the study of both physical and chemical systems in equilibrium. It provides the student with instruction that will enable him to predict products in solubility, acid-base, and oxidation-reduction reactions and to write and balance equations for these reactions and solve problems involving equilibria constants. Students will gain more from this course if they have had some previous work in energy concept of atoms and molecules. The booklet lists the relevant state-adopted texts and states the performance objectives for the unit. It provides a course outline; suggests experiments, demonstrations, projects, reports, speakers, and field trips; and presents a series of related mathematical problems. Films, film loops, filmstrips, and several reference books are recommended. A master sheet is provided relating each suggested activity to the specific performance objectives. (EB)

ED 093655

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AUTHORIZED COURSE OF INSTRUCTION FOR THE QUINMESTER PROGRAM

DYNAMIC NATURE OF ATOMS AND MOLECULES

5316.06

SCIENCE

(Experimental)

DADE COUNTY PUBLIC SCHOOLS

DIVISION OF INSTRUCTION•1971

SE 017 612

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DYNAMIC NATURE OF ATOMS AND MOLECULES

5316.06

SCIENCE

(Experimental)

Written by Jacquelin F. Buffaloe
for the
DIVISION OF INSTRUCTION
Dade County Public Schools
Miami, Florida
1972

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COURSE DESCRIPTION

In the course, Dynamic Nature of Atoms and Molecules, students will study material which will enable them to (1) recognize both physical and chemical systems in equilibrium (2) predict products in solubility, acid-base, and oxidation-reduction reactions (3) write and balance equations for the reactions mentioned and (4) solve problems involving equilibria constants.

ENROLLMENT GUIDELINES

This quin is adaptable for Nursing Chemistry and all other first year chemistry classes on this subject. Students should have successfully completed Introduction to Chemistry, Reactions of Atoms and Molecules, and Scientific Mathematics or shown their readiness by passing a test. Students will gain a more through understanding of equilibrium from a study of this quin if they have also been successful in Energy of Atoms and Molecules.

STATE ADOPTED TEXTS

1. Choppin, Gregory R. and Jaffe, Bernard. *Chemistry: Science of Matter, Energy and Change*. Morristown, New Jersey: Silver Burdett Company, 1966.
2. Greenstone, Arthur W.; Sutman, Frank S.; and Hollingworth, Leland G. *Concepts in Chemistry*. New York: Harcourt, Brace and World, Inc., 1966.
3. Metcalfe, Clark R.; Williams, John E.; and Castka, Joseph F. *Modern Chemistry*. New York: Holt, Rinehart and Winston, Inc., 1966.
4. O'Connor, Paul R.; Davis, Joseph E. Jr.; Haenisch, Edward L.; MacNab, W. Keith; and McClellan, A. L. *Chemistry: Experiments and Principles.* Lexington, Mass.: Raytheon Education Company, 1968.

PERFORMANCE OBJECTIVES

1. Given a set of descriptions of various types of systems, the student will in each case state those additional properties which are needed to create an equilibrium.

2. Given an equation for a chemical system at equilibrium which shows the heat involved and the state of each chemical, the student will state the reaction and chemicals favored when each change listed below occurs.

   (1) temperature increase
   (2) addition of a catalyst
   (3) pressure decrease
   (4) volume decrease
   (5) increase in the concentration of one of the reactants

3. Given a list of equations of chemical systems at equilibrium involving solids, liquids, gases, and solvents, the student will write equilibrium constant expressions.

4. Given three equations representing systems at equilibrium and their equilibrium constants at any constant temperature for all systems, the student will place them in order of degree of reaction to the right (highest degree first).

5. Given a chemical equation of a system at equilibrium, the initial concentrations of each chemical involved, and the equilibrium concentration of one product, the student will determine the equilibrium constant.

6. Given the equation of a system at equilibrium, the equilibrium constant, and the initial concentration of each chemical, the student will calculate the equilibrium concentration of each reactant and product. (NOTE TO INSTRUCTORS: Select problems which avoid lengthy arithmetic calculations and require the extraction of no root other than the square root.)

7. Given the formulas of two solutes in separate solutions, the concentrations of the solutions, the relative volumes of solutions to be mixed, a table of solubilities, and a table of Ksp values, the student will determine if a precipitate will form.

8. Given the formula and the Ksp value for a chemical, the student will calculate the solubility of the chemical.

PERFORMANCE OBJECTIVES (Continued)

9. Given the formulas of two chemicals and a table showing the relative strengths of acids, the student will write a balanced equation for an acid-base reaction where the acid defined by the Lowry-Brønsted Theory of acids and bases.

10. Given an equation for an acid-base reaction and a table of relative acid strengths, the student will state the chemicals favored at equilibrium.

11. Given five (5) sets of reactants which will undergo oxidation-reduction reactions, and a table of oxidation half-reactions and their standard potentials, the student will select those sets of reactants which will react spontaneously.

12. Given a table of oxidation half-reactions with their standard potentials, the reactants to be used and a specific product to be made, the student will draw and label an electrical cell using words, chemical formulas, arrows, or equations in the appropriate places on the drawing to show:

    (1) essential parts of the cell,
    (2) composition of the solutions and electrodes,
    (3) direction of flow of the electrons,
    (4) half-reactions occuring in each part of the cell,
    (5) oxidation half-cell, and
    (6) the oxidizing agent.

13. Given a list of five sets of reactants and a table of oxidation half-reactions with their standard potentials, the student will write a balanced equation for each set of reactants.

14. Given a list of five oxidation-reduction equations and a table of half-reactions with their standard potentials, the student will calculate the voltage which would be produced for each reaction if it were taking place in an electrical cell.

15. Given the formulas for the reactants and products in oxidation-reduction reactions, the student will balance the equations using both the half-reaction and the oxidation number methods.

COURSE OUTLINE

I. Properties and Factors Which Determine Equilibrium

  A. Vapor-liquid equilibrium

  B. Systems at equilibrium at the boiling and melting points

  C. Solubility equilibrium

  D. Chemical equilibrium

II. Factors Which Affect the State of Equilibrium (LeChatelier's Principle)

  A. Temperature changes

  B. Concentration changes

  C. Addition of a catalyst

  D. Pressure changes

  E. Volume changes

III. Introduction to the Equilibrium Expression and Constant

  A. Basis for the derivation of the general expression - ratio

  B. How to write the equilibrium expression

  C. Determination of the K value

  D. Significance of a large or small value for K

IV. Determination of the Concentration of Various Particles in a Chemical Equilibrium by Using the Equilibrium Constant

V. Types of Chemical Reactions Involving Equilibria Situations

  A. Precipitation

    1. Writing the Ksp expressions and determination of the Ksp value

    2. Use of the Ksp in the determination of

      (a) Formation of a precipitate
      (b) Solubility of a chemical
      (c) Concentration of an ion needed to precipitate a compound

B. Acid-base Reactions

1. Strength of acids and bases - use Ka to determine relative strength

2. Writing equations for reactions between strong acids and bases and weak acids and bases

3. Use of the Ka to determine the chemicals favored at equilibrium

4. Determination of the pH of a weak acid by using Ka to determine the hydrogen ion concentration

5. Introduction of Kw and its use in the determination of the hydrogen and hydroxide ion concentrations in aqueous solutions

C. Oxidation - Reduction Reactions

1. Definition of oxidation, reduction, oxidizing agent, and reducing agent

2. Elementary definition of electricity and its production

3. Voltage

   (a) Definition
   (b) Calculation

4. Description of an electrical cell

   (a) Parts
   (b) Function

5. Electrical cell at equilibrium - voltage

6. Balancing oxidation - reduction equations

   (a) Oxidation number method
   (b) Half-reaction method

EXPERIMENTS

Barrett, Richard L. and Price, Jack. Chemistry, A Modern Course Laboratory Manual. Columbus, Ohio: Charles E. Merrill Books, Inc., 1965.

1. Equilibrium (exp. 33, p.71)
2. Measurement of pH by Indicators (exp. 34, p.73)
3. Oxidation-reduction Reactions (exp. 29, p.59)
4. Precipitation (exp. 35, p.77)
5. Titration (exp. 30, p.63)

Castka, Joseph F.; Metcalfe, H. Clark; and Williams, John E. Exercises and Experiments in Chemistry. New York: Holt, Rinehart and Winston, 1966.

6. Brønsted Acid and Bases - Indicators (exp. 28, p.195)
7. Equilibrium (exp. 39, p.235)
8. Hydronium Ion Concentration, pH (exp. 27, p.193)
9. Oxidation - Reduction Reactions (exp. 43, p.243)
10. Preparation and Properties of Acids (exp. 25, p.185)
11. Preparation and Properties of Hydroxides (exp. 26, p. 189)
12. Relative Strengths of Acids and Bases (exp. 29, p.199)
13. Solubility Product Constant (exp. 40, p.239)
14. Titration of Acid and a Hydroxide (exp. 30, p.201)

Chemistry 2: Advanced Placement Course Outline and Manual of Laboratory Activities. Curriculum Bulletin 8-D. Dade County Public Schools, 1966.

15. Electrochemical Cells (exp. 31, p.147)
16. Heat of Neutralization (exp. 26, p.129)
17. Hydronium Ion Concentration (exp. 11, p.84)
18. Relative Strengths of Acids and Bases (exp. 13, p.91)
19. Solubility (exp. 21, p.113)
20. Solubility Product Constants (exp. 22, p.116)
21. Titration of Acid and Hydroxide (exp. 14, p.93)

Davis, Joseph E.; MacNab, W. Keith; Haenishch, Edward L.; McClellan, A.L.; and O'Connor, Paul R. Laboratory Manual for Chemistry Experiments and Principles. Lexington, Mass.: Raytheon Education Company, 1968.

22. An Introduction to Oxidation - Reduction (exp. 34, p.87)
23. Applying LeChatelier's Principle (exp. 30, p.76)
24. Chemical Equilibrium (exp. 29, p.73)
25. Determination of $[H^+]$ Using Indicators (exp. 33, p.84)
26. Heat of Acid-Base Reaction (exp. 31, p.78)
27. Reactions of Ions in Aqueous solution (exp. 17, p.50)

EXPERIMENTS (Continued)

Ferguson, Harold; Schmuckler, Joseph; and Siegelman, Irwin. Investigating Matter, Energy and Change. Morristown, New Jersey: Silver Burdett Company, 1966.

28. Acid-Base Indicators (exp. 23, p.169)
29. Acid-Base Titration (exp. 24, p.179)
30. An Introduction to Reduction and Oxidation Potential (exp. 25, p.185)
31. Chemical Equilibrium (exp. 19, p.145)
32. Hydrogen as a Reducing Agent (exp. 15, p.113)
33. Solutions and Solubility (exp. 20, p.153)
34. Solution Species and Precipitates (exp. 21, p.157)
35. Solubility and Solubility Products (exp. 22, p.163)

Geffner, Saul and Lauren, Paul. Experimental Chemistry and Workbook. New York: Amsco Publications, 1968.

36. Determination of a Solubility Product Constant of a Salt (exp. 23, p.91)
37. Determination of the Concentration of a Basic Solution (exp. 27, p.107)
38. Determination of the Concentration of an Acid Solution (exp. 28, p.111)
39. Energy and Chemical Change (exp. 16, p.65)
40. Equilibrium (exp. 21, p.85)
41. Electrolysis (exp. 33, p.131)
42. Ionic Equilibrium (Lechatelier's Principle) (exp. 22, p.87)
43. Making a Solubility Curve (exp. 9, p.37)
44. Nonreversible Reactions (exp. 29, p.113)
45. Oxidation Potentials and Electron Transfer (exp. 32, p.125)
46. pH and Indicators (exp. 24, p.95)
47. Principles and Properties of Acids (exp. 25, p.99)
48. Principles and Properties of Bases (exp. 26, p.103)
49. Solutions (exp. 8, p.33)

Greenstone, Arthur W. Concepts in Chemistry Laboratory Manual. Atlanta: Harcourt, Brace and World, 1966.

50. Acids, Bases, and Salts (exp. 21, p.74)
51. Chemical Equilibrium Reversible and Irreversible Reactions (exp. 25, p.83)
52. Electrochemistry (exp. 32, p.100)
53. Indicators (exp. 24, p.81)
54. Ionization Constant of a Weak Acid or Base (exp. 27, p.88)
55. Preparation of Acids and Bases (exp. 22, p.76)
56. Reduction by Hydrogen (exp. 7, p.36)
57. Titration (exp. 23, p.78)
58. Voltaic Cells (exp. 31, p.97)

EXPERIMENTS (Continued)

Pimentel, George C. ed. Chemistry: An Experimental Science Laboratory
Manual. San Francisco: W. H. Freeman and Company, 1963.

59. An Introduction to Oxidation-Reduction (exp. 20, p.55)
60. Applying LeChatelier's Principle to Some Reversible Chemical Reactions (exp. 19, p.53)
61. Chemical Equilibrium (exp. 15, p.43)
62. Determination of $[H^+]$ of Solutions Using Indicators (exp. 18, p.50)
63. Determination of Ksp of $AgCH_3COO$ (exp. 16, p.46)
64. Electrochemical Cells (exp. 21, p.57)
65. Heat of Some Acid-Base Reactions (exp. 17, p.48)
66. Melting Temperature of a Pure Substance (exp. 3, p.4)
67. Quantitative Titration (exp. 23, p.61)

Smith, William T., Jr. and Wood, Jesse H. Laboratory Manual for College
Chemistry. New York: Harper and Row, Publishers, 1966.

68. Chemical Equilibria (exp. 24, p.101)
69. Electrochemistry (exp. 30, p.127)
70. Solubility Product Constant (exp. 25, p.107)

Weaver, Albert C. Scientific Experiments in Chemistry. New York: Holt,
Rinehart, and Winston, Inc., 1966.

71. Acid and Alkaline Solutions (exp. 27, p.52)
72. Elementary Volumetric Analysis (Acid-base titration) (exp. 30, p.57)
73. Neutralizing Phosphoric Acid (exp. 31, p.59)
74. Oxidation-Reduction (exp. 32, p.62)
75. Progressive Precipitation (exp. 21, p.38)
75. Reactions Between Oxides and Water (exp. 25, p.47)
77. Reactions of Lead Chloride and Aluminum (exp. 46, p.86)
78. Reaction of Zinc and Nitric Acid (exp. 49, p.94)
79. Solubility Product Constant of Cadinum Hydroxide (exp. 45, p.86)

DEMONSTRATIONS

Alyea, Hubert N. and Dutton, Frederic B. Tested Demonstrations in General Chemistry. Easton, Penn.: Division of Chemical Education of the American Chemical Society, 1966.

1. Vapor Pressure of Pure Liquids (p. 65)
2. Solubility phase Equilibrium at Room Temperature (p. 67)
3. Electric Cells (p. 79)
4. Reversible Reactions - Shifting Equilibrium - Models Illustration Chemical Equilibrium Effect of Concentration, effect of Temperature, Common Ion Effect, Ksp. (pp. 86, 87)
5. Oxidation and Reduction (pp. 87, 88, 89)
6. Introduction to LeChatelier's Principle (p. 194)
7. A Reaction to Introduce Lewis Acid-Base Chemistry (p. 194)
8. Indicators and pH (p. 147)
9. Electrode Potentials - Hydrogen Electrode (p. 141)

PROJECTS

1. Prepare a colorimeter.

2. Determine the range of pH over which a natural indicator such as cabbage or grape juice could be used.

3. Devise a general method to determine the Ksp and check its variation with temperature, amount of solid present, etc..

4. Prepare a solution of a weak acid or base, or the salt of such an acid or base. From the pH determine the equilibrium constant. Test various factors and their effect on pH and the equilibrium constant.

5. Design an electroplating cell.

6. Produce elemental materials by electrolysis.

7. Make a voltaic or galvanic cell having a gas or other unusual electrode. Determine standard potentials.

REPORTS

1. Choose equations for various types of reactions and states of materials and describe changes that would result if temperature, pressure, concentration, etc. were changed. Explain each.
2. Describe several situations in nature and the industrial area where chemical equilibria become important.
3. How does a fuel cell work?
4. Describe the types of calculations that can be made from an equation and an equilibrium constant. Give examples.
5. Give examples and explain (based on structure and atomic properties) the trends in the strength of binary and ternary acids of any chemical family or period.
6. Explain the action and selection of indicators.
7. Research a variety of types of catalysts and describe the mechanisms by which they work.
8. Explain why some salt solutions are acidic and others are not. Examples and methods to be used in predictions should be included.
9. A better student might report on variations of voltage and concentration of the solutions.
10. How to use a pH meter - principle behind operation.

SPEAKERS AND FIELD TRIPS

1. City of Miami Water Plants
   Director of Dept. of Water and Sewers
   665-7471

2. Dade County Air and Water Pollution Control
   864 N.W. 23rd Street
   377-5891

3. University of Miami School of Medicine
   Director of Public Relations
   Sewell Building
   1477 N.W. 12th Avenue
   350-6256

RELATED MATH PROBLEMS

1. To produce a saturated solution of calcium fluoride, CaF_2, you need to dissolve 0.0068 grams of CaF_2 per 0.250 liters of solution. What is the Ksp of CaF_2?

2. It takes 0.048 grams of BaF_2 to saturate 15.0 ml of water (consider this total volume). What is the Ksp of BaF_2?

3. Given that the Ksp of $BaSO_4$ is 1.5×10^{-9}, how many grams of barium sulfate can you dissolve in 1000 liters of solution?

4. Given that the Ksp of Thallous Iodide, TlI, is 8.9×10^{-8}, calculate the grams of TlI that can be dissolved in 1.25 liters of solution.

5. Given that the Ksp of $BaSO_4$ is 1.5×10^{-9}, would you expect a precipitate to form if 10.0 ml of 0.0100 M $BaCl_2$ solution were mixed with 30.0 ml of 0.0050 M Na_2SO_4 solution?

6. The Ksp of $SrSO_4$ is 7.6×10^{-7}. Should precipitation occur when 25.0 ml of 1.0×10^{-3} M $SrCl_2$ solution is mixed with 15.0 ml of 2.0 x 10 - 3 M Na_2SO_4?

7. You are given a box containing NH_3, N_2, and H_2 at equilibrium at 1000°K. Analysis shows that the concentration of NH_3 is 0.102 M, N_2 is 1.03 M and H_2 is 1.62 M. Calculate K for the reaction

$$2NH_3\ (g) \rightleftarrows N_2\ (g) + 3\ H_2\ (g)$$

RELATED MATH PROBLEMS (Continued)

8. Hydrogen gas, sulfur vapor and hydrogen sulfide gas are in equilibrium with each other under the following condition: 1.68 moles of hydrogen gas, 2.88×10^{-5} moles of sulfur vapor and 1.37 moles of hydrogen sulfide gas are in a volume of 18.0 liters at 750°C. Calculate the K for the reaction.

9. You are given that K is 2.37×10^{-3} for $N_2(g) + 3H_2(g) \rightleftarrows 2\ NH_3(g)$ at 1000°K. If the concentrations of N_2 is fixed at 2.00M and that of H_2 at 3.00 molar and the equilibrium is to remain, what must the concentration of NH_3 be?

10. Given K is 0.0224 for $PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$ at 500°K. What must the equilibrium concentration of Cl_2 be in this system when the concentration of PCl_5 is 0.0428 M and that of PCl_3 is 1.32 M?

11. Given a solution which is labeled 0.0010 M HNO_2, what is the concentration of H^+ and NO_2, in this solution? Consider the part of HNO_2 which ionized to be insignificant when compared to the concentration of the acid. The Ka of HNO_2 is 4.5×10^{-4}.

12. If you dissolve 1.75 grams of HClO in enough water to make 500 ml of solution, what would be the concentrations of H^+, ClO^-, and HClO in the final solution? $Ka = 3.2 \times 10^{-8}$. Ignore the amount of HClO that ionizes when determining the concentration of that chemical.

13. You have a solution that is 0.220 M HNO_2 and the Ka is 4.55×10^{-4}. Calculate the $[H^+]$ concentration and from this determine the pH of the solution.

14. The Ka for $HCNO \rightleftarrows H^+ + CNO^-$ is 1.2×10^{-4}. If you want the solution to have a pH of 3, what molarity of solute do you need?

15. When 1 mole of pure chemical X is mixed with one mole of pure chemical Y at room temperature, the equilibrium mixture contains 2/3 mole of each product, chemical Z and water. What is the equilibrium constant? These chemicals are mixable and water is not considered the solvent.

$$X + Y \rightleftarrows Z + H_2O$$

16. The $[H^+]$ in a 0.1 M solution of benzoic acid ($HC_7H_5O_2$) is 2.6×10^{-3} M. Calculate the Ka of the acid.

17. The pH of a solution is 5. Calculate the $[OH^-]$.

RELATED MATH PROBLEMS (Continued)

18. Balance the following equations using both the oxidation number and half-reaction methods.

(1) $Zn + H^+ + NO_3^- \longrightarrow Zn^{+2} + H_2O + NO$

(2) $I_2 + S_2O_3^{-2} \longrightarrow S_4O_6^{-2} + I^-$

(3) $SbH_3 + Hg_2 \longrightarrow Sb + Hg + H^+$

(4) $MnO + PbO_2 + H^+ \longrightarrow MnO_4^- + Pb^{+2} + H_2O$

(5) $CdS + I_2 \longrightarrow Cd^{+2} + I^- + S$

19. Determine the voltage for the following reactions.

(1) $Zn + Cu^{+2} \longrightarrow Cu + Zn^{+2}$

(2) $H_2 + 2OH^- + I_2 \longrightarrow 2\ I^- + 2\ H_2O$

(3) $3\ Fe^{+2} + Au^{+3} \longrightarrow 3\ Fe^{+3} + Au$

(4) $SO_2 + 2\ H_2O + 2\ Ag^+ \longrightarrow SO_4^{-2} + 4\ H^+\ \ 2\ Ag$

(5) $3Mn^{+2} + Cr_2O_7^{-2} + 14\ H^+ \longrightarrow 3\ MnO_2 + 12\ H^+ + 2\ Cr^{+3} + H_2O$

FILMS

Available From Dade County Audiovisual Center

1. Acids, Bases, and Salts
   1-10947, 21', C
2. Acid-Base Indicators
   1-10799, 19', C
3. Catalysis
   1-10809, 16', C
4. Catalysis
   1-01959, 11', B/W
5. Electrochemical Cells
   1-10834, 22', C
6. Equilibrium
   1-10829, 22', C
7. Oxidation and Reduction
   1-01934, 11', B/W
8. Ionic Equilibrium
   1-10930, 16', C
9. Properties of Acids, Bases and Salts
   1-30347, 28', C
10. Standard Solutions and Titration
    1-10926, 21', B/W

Available From Encyclopaedia Britannica Films Inc., 1150 Wilmette Avenue, Wilmette, Illinois

11. Electrochemistry
    11', B/W

Available From Yale Chemistry Films, Association Films Inc., 347 Madison Avenue, New York, New York 10017

12. LeChatelier's Principle
    YF225, 3', C
13. Solubility Product
    YF237, 7', C

FILM LOOPS

Kalmia Company, Concord, Mass. 01742

1. Titration
   5400, C

Ealing Film-Loops, 2225 Massachusetts Avenue, Cambridge, Mass. 02140

2. Boiling Points and Pressure
   80-3403

3. Equilibrium: LeChatelier's Principle
   84-0223

4. Equilibrium: Solutions and Precipitates
   84-0215

5. Galvanic Cells: Half-Cell Reactions
   84-0207

6. Galvanic Cells: Electrolytic and Electrodes
   84-0199

7. Oxidation and Reduction: Electrolytic Cells
   84-0157

8. Oxidation and Reduction : The Halogens
   84-01401

FILM STRIPS

McGraw-Hill Book Co., Text Film Dept., 330 W. 42nd St., N.Y., N.Y.

1. Acid and Basic Solutions
2. Balancing Equation (Oxidation-Reduction)
3. Chemical Bond
4. Electrolysis

FILM STRIPS (Continued)

5. Equilibrium
6. Ionization
7. Solubility Product Constant

Film Strip of the Month Club, 355 Lexington Ave., N.Y. 17, N.Y.

8. Acid-Base Theories
9. Equilibrium Constant
10. The Mole Concept
11. Oxidation-Reduction

SUGGESTED DISCUSSION QUESTIONS

1. What are the characteristics that distinguish a system at equilibrium from a "steady state?"
2. Explain "dynamic nature" of an equilibrium.
3. How do you know a chemical has reached the melting or boiling point?
4. Would a pressure change affect all systems at equilibrium? Explain, give an example.
5. How can the equilibrium constant be used?
6. How are K, Ka, Ksp, Kw etc. similiar? How are they different?
7. Why does solubility vary?
8. Can you make all chemical solutions precipitate? Explain your answer.
9. What factors affect a system at equilibrium? Explain why in each case.
10. Use LeChatelier's Principle to explain what happens in an equilibrium when the concentrations are increased.

SUGGESTED DISCUSSION QUESTIONS (Continued)

11. How does the structure of an acid or base determine its strength?

12. What is the difference in the strength and the concentration of an acid or base? Give an example.

13. Why is it necessary for a galvanic or voltaic cell to have two containers and a salt bridge? Would oxidation-reduction reactions occur without them?

14. What rules should you follow in setting up a galvanic or voltaic cell?

15. What makes the electrons flow through the wire in a galvanic or voltaic cell?

16. What rules should you follow to determine the products in a reaction between a weak acid and a weak base? How do you tell the chemicals or reaction favored?

17. How do you use a table of half-cell oxidation reactions and potentials to (1) determine the spontaneous reactions and (2) complete equations?

18. What rules must you use to write the expression for an equilibrium constant? Explain your answer.

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